



CIE A Level Chemistry



Your notes

6.2 Properties of Transition Elements (A Level Only)

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- * 6.2.2 Oxidation States of Transition Metals
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- * 6.2.6 Geometry of Complexes
- * 6.2.7 Degenerate & non-Degenerate d Orbitals
- * 6.2.8 Coloured Complexes

Your notes

		TRANSITION METALS																0
1	2											3	4	5	6	7	He	
Li	Be											B	C	N	O	F	Ne	
Na	Mg											Al	Si	P	S	Cl	Ar	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
Fr	Ra	Ac																

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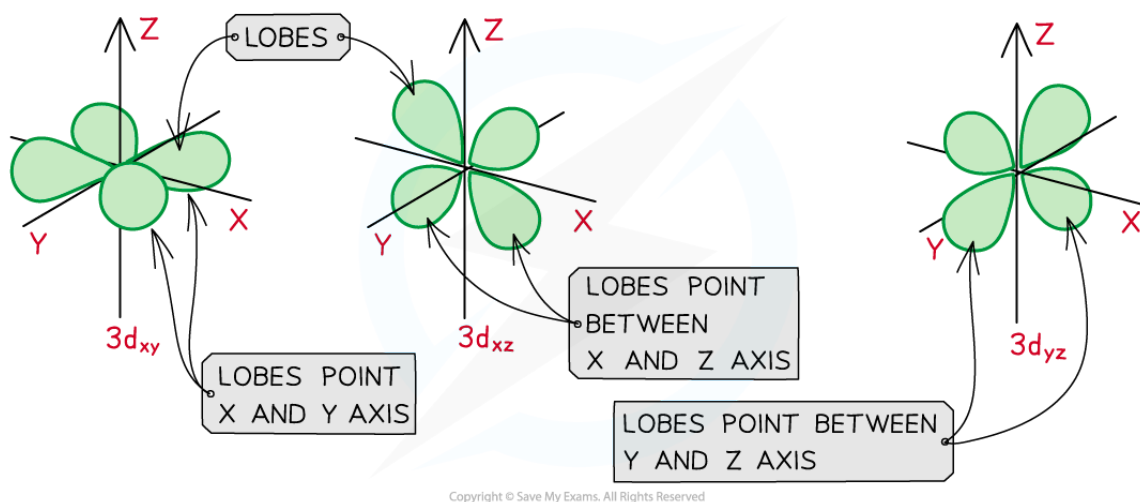
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Shape of $3d(xy)$ & $3d(z^2)$ Orbitals

- The transition elements all have **incomplete** d subshells
- There are five orbitals in a d subshell. Some of these orbitals may have similar **shapes** but different **orientations**, whereas others may have completely different **shapes**
- The five orbitals are
 - $3d_{yz}$
 - $3d_{xz}$
 - $3d_{xy}$
 - $3d_{x^2-y^2}$
 - $3d_{z^2}$
- Note that students are required to sketch the shapes of the $3d_{xy}$ and $3d_{z^2}$ orbitals **only**

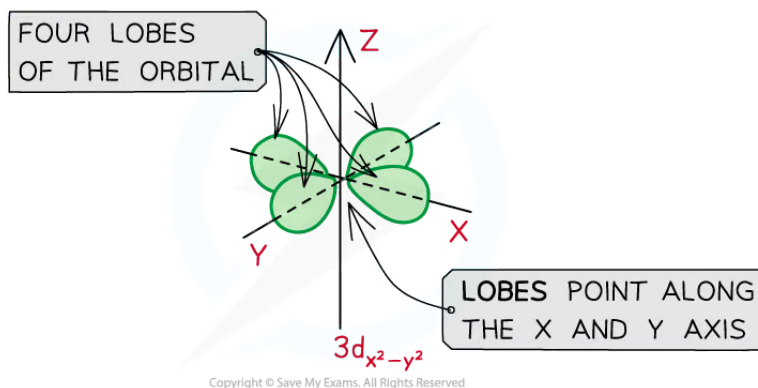
Shapes of the $3d$ orbitals

- The $3d_{yz}$, $3d_{xz}$, and $3d_{xy}$ orbitals are orbitals which lie in the y-z, x-z and x-y plane respectively
 - They all have **four lobes** that point **between** the two axes



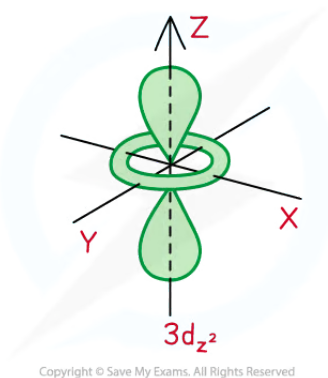
The $3d_{yz}$, $3d_{xz}$, and $3d_{xy}$ orbitals all have four lobes which are similar in shape but point between different axes

- The $3d_{x^2-y^2}$ orbital looks like the $3d_{yz}$, $3d_{xz}$, and $3d_{xy}$ orbitals, as it also consists of **four lobes**
- The difference is that these lobes point **along** the x and y axes and **not** between them



The four lobes in a $3d_{x^2-y^2}$ orbital point along the axes

- The $3d_{z^2}$ orbital is different from the other orbitals, as there are two main lobes which form a **dumbbell shape**
- The two main lobes point **along** the z-axis and there is a “doughnut” ring around the centre



The $3d_{z^2}$ orbital has a dumbbell shape with a ring around the centre



Your notes

Properties of the Transition Elements

- Although the **transition elements** are metals, they have some properties unlike those of other metals on the periodic table, such as:
 - Variable **oxidation states**
 - Behave as **catalysts**
 - Form **complex ions**
 - Form **coloured compounds**

Ions of transition metals

- Like other metals on the periodic table, the transition elements will lose electrons to form positively charged ions
- However, unlike other metals, transition elements can form more than one positive ion
 - They are said to have **variable oxidation states**
- Because of this, Roman numerals are used to indicate the oxidation state on the metal ion
 - For example, the metal sodium (Na) will only form Na^+ ions (no Roman numerals are needed, as the ion formed by Na will always have an oxidation state of +1)
 - The transition metal iron (Fe) can form Fe^{2+} (Fe(II)) **and** Fe^{3+} (Fe(III)) ions
- The table below shows the most common oxidation states of a few transition metals

Oxidation states of transition elements table



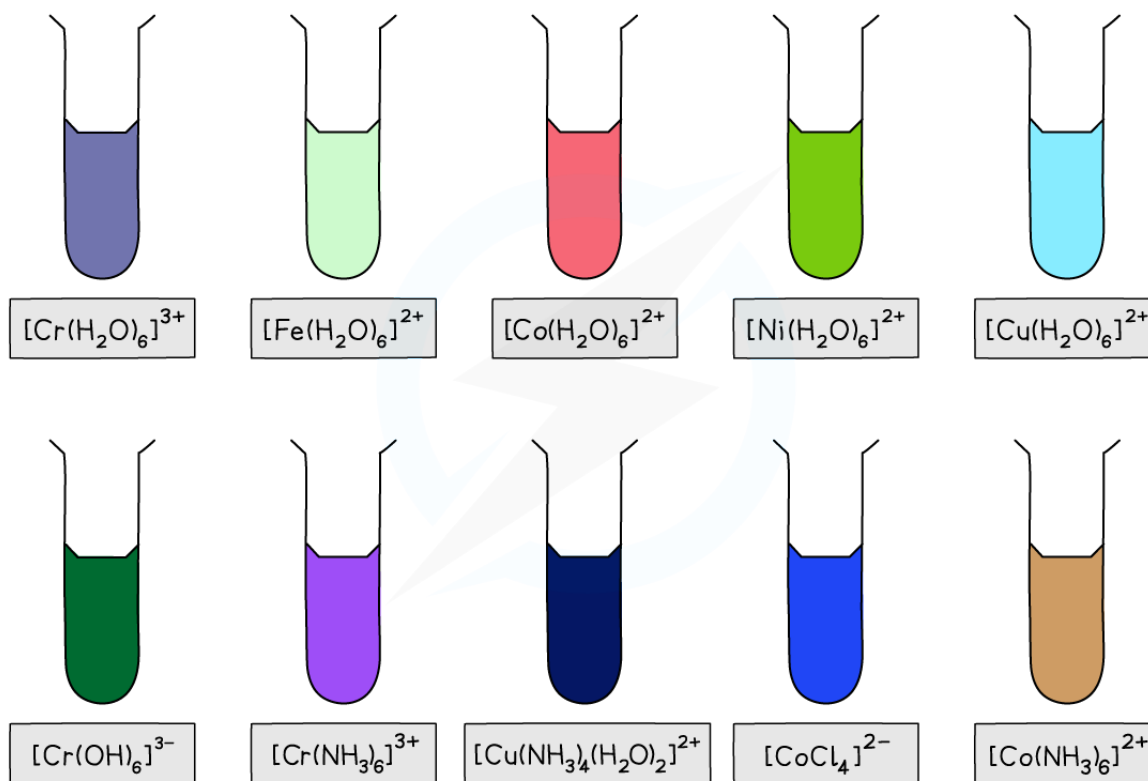
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Transition Element	Common oxidation States
Ti	+3, +4
V	+2, +3, +4, +5
Cr	+3, +6
Mn	+2, +4, +6, +7
Fe	+2, +3
Ni	+2
Cu	+1, +2

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Coloured complex

- Another characteristic property of transition elements is that their compounds are often **coloured**
 - For example, the colour of the $[\text{Cr}(\text{OH})_6]^{3-}$ complex (where oxidation state of Cr is +3) is **dark green**
 - Whereas the colour of the $[\text{Cr}(\text{NH}_3)_6]^{3+}$ complex (oxidation state of Cr is still +3) is **purple**



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Examples of some transition metal ions and their coloured complexes

Transition elements as catalysts

- Since transition elements can have variable oxidation states, they make excellent **catalysts**
- During catalysis, the transition element can change to various oxidation states by gaining electrons or donating electrons from reagents within the reaction
 - For example, iron (Fe) is commonly used as a catalyst in the Haber Process, switching between the +2 and +3 oxidation states
- Substances can also be adsorbed onto their surface and activated in the process

Complex ions

- Another property of transition elements caused by their ability to form variable oxidation states, is their ability to form **complex ions**
- A complex ion is a molecule or ion, consisting of a central metal atom or ion, with a number of molecules or ions surrounding it
- The molecules or ions surrounding the central metal atom or ion are called **ligands**

- Due to the different oxidation states of the central metal ions, a different number and wide variety of ligands can form bonds with the transition element
 - For example, the chromium(III) ion can form $[\text{Cr}(\text{NH}_3)_6]^{3+}$, $[\text{Cr}(\text{OH})_6]^{3-}$ and $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$ complex ions



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6.2.2 Oxidation States of Transition Metals

Effects of the 3d & 4s Subshells on Oxidation States of the Transition Elements

- Transition elements can have **variable oxidation states**
- These variable oxidation states can be formed as the 3d and 4s atomic orbitals are **similar in energy**
- This means that a similar amount of energy is needed to remove a different number of electrons
- When the transition elements form **ions**, the electrons of the 4s subshell are lost first, followed by the 3d electrons
- The most common oxidation state is +2, which is usually formed when the two 4s electrons are lost

Oxidation number at the start of the 3d transition elements

- At the **start** of the period, it is easier for the transition elements to lose the **maximum** number of electrons
- The maximum oxidation number of these transition elements involves all the 4s and 3d electrons in the atom
- For example, the maximum oxidation state of a titanium (Ti) ion is +3 or +4, as two 4s electrons and either 1 or 2 3d electrons are lost
- Ti atom = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$
- Ti³⁺ ion = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$
 - Ti⁴⁺ ion = $1s^2 2s^2 2p^6 3s^2 3p^6$

Oxidation number at the end of the 3d transition elements

- Towards the **end**, the 3d transition elements are more likely to adopt the +2 oxidation state
- This is because across the d block, the 3d electrons become slightly **harder** to remove as the **nuclear charge** increases
 - The 3d electrons are attracted more strongly to the nucleus
 - The higher oxidation states become less stable
- Therefore, the elements are more likely to lose their 4s electrons only
- For example, nickel (Ni) is a transition element at the end of the period which only forms ions with oxidation state +2, due to loss of the 4s electrons only
 - Ni atom = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$
 - Ni²⁺ ion = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$



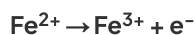
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Transition Elements: Catalysts

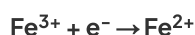
- Transition elements are often used as **catalysts** due to their ability to form ions with more than one **stable oxidation state**, and the fact that they contain **vacant d orbitals**

Oxidation states

- Transition element ions can adopt more than one **stable oxidation state**
- This means that they can accept and lose electrons easily to go from one oxidation state to another
- They can therefore catalyse **redox** reactions, by acting as both **oxidising agents** and **reducing agents**
- For example, iron (Fe) is often used as a catalyst due to its ability to form Fe(II) and Fe(III) ions, acting as an oxidising agent and a reducing agent
 - When Fe(II) acts as a reducing agent, it will reduce another species and become oxidised itself



- The Fe^{3+} formed in the catalytic cycle, can then also act as an oxidising agent by oxidising another species and getting reduced itself to reform the Fe^{2+} ion



- Transition element ions with high oxidation states make powerful oxidising agents, because they will readily accept electrons
 - A common example of this is potassium permanganate (VII), where manganese has an oxidation state of +7

Vacant d orbitals

- When transition elements form ions, they have **vacant d orbitals** which are **energetically accessible**
 - The orbitals are not too high in energy
- This means that **dative bonds** can be formed between the transition element ion and **ligands**
 - Each ligand provides the pair of electrons required for the formation of a bond between the ion and the ligand
 - This pair of electrons is donated into the ion's vacant d orbital
- The table below shows the electron configuration of the transition element **atoms**
- When they form ions, empty **d orbitals** are obtained which can be filled by the pairs of electrons donated by the ligands

Electronic configuration of transition elements table



Your notes

Element	Electronic Configuration
Ti	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$
V	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$
Cr	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$
Mn	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
Fe	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$
Co	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$
Ni	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$
Cu	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

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Transition Metals: Complex Ions

- A **complex** is a molecule or ion formed by a central metal atom or ion surrounded by one or more **ligands**
 - A complex can have an overall positive or negative charge, or it can be neutral
 - If a complex is charged overall, it is often called a **complex ion**
- Transition elements can easily form complex ions, because they have **empty d orbitals** that are energetically accessible
 - The empty d orbitals are therefore not too high in energy and can accommodate a lone pair of electrons
- The transition element in the centre will accept pairs of electrons from the ligands into their empty d orbitals, forming **dative bonds**
 - The transition element in the centre is often referred to as the **central metal ion**, as all transition elements are metals, and it is often an ion in the centre
- For example, the titanium(III) (Ti^{3+}) ion, has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$
 - This means that there are vacant d orbitals that can be occupied by electrons, from ligands such as H_2O for example, to form a $[\text{Ti}(\text{H}_2\text{O})_6]^{3+}$ complex ion
 - 6 water ligands have each donated a pair of electrons, to form 6 dative bonds with the central metal ion



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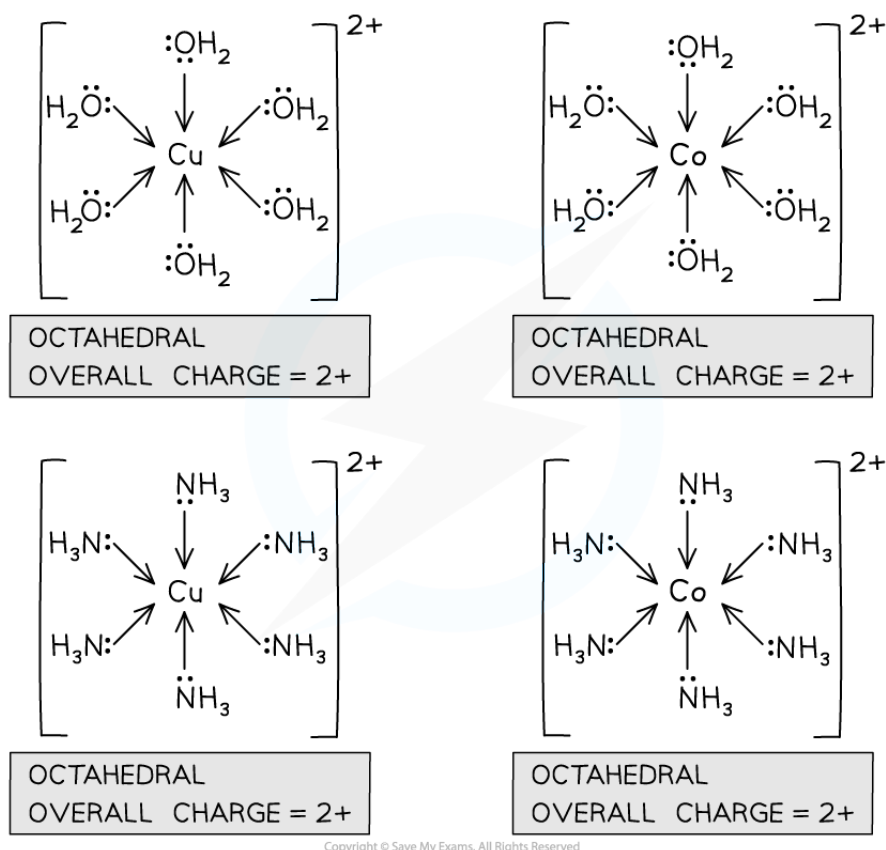
6.2.3 Transition Metal Complexes

Transition Elements: Ligands & Complex Formation

- Transition element ions can form complexes which consist of a **central metal ion** and **ligands**
- Copper(II) and cobalt(II) ions will be used as examples of the central metal ions, in the complex formation with water (H_2O), ammonia (NH_3), hydroxide (OH^-), and chloride (Cl^-) ligands
 - A copper(II) ion has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9$
 - A cobalt(II) ion has an electronic configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7$

Complexes with water & ammonia molecules

- Water and ammonia molecules are examples of **neutral** ligands
- Both ligands contain a lone pair of electrons which can be used to form a dative covalent bond with the central metal ion
 - In water, this is the lone pair on the oxygen atom
 - In ammonia, it is the lone pair on the nitrogen atom
- Since water and ammonia are small ligands, 6 of them can fit around a central metal ion, each donating a lone pair of electrons, forming 6 dative bonds
 - The **coordination number** of a complex is the number of dative bonds formed between the central metal ion and the ligands
 - Since there are 6 dative bonds, the coordination number for the complex is 6
- Complexes with a coordination number of 6 have an **octahedral** shape
- The overall charge of a complex is the sum of the charge on the central metal ion, and the charges on each of the ligands
- A complex with cobalt(II) or copper(II) as a central metal ion, and water or ammonia molecules as ligands, will have an **overall charge** of **2+**
 - The central metal ion has a 2+ charge and the ligands are neutral



Cobalt(II) and copper(II) form octahedral complexes with ammonia and water ligands

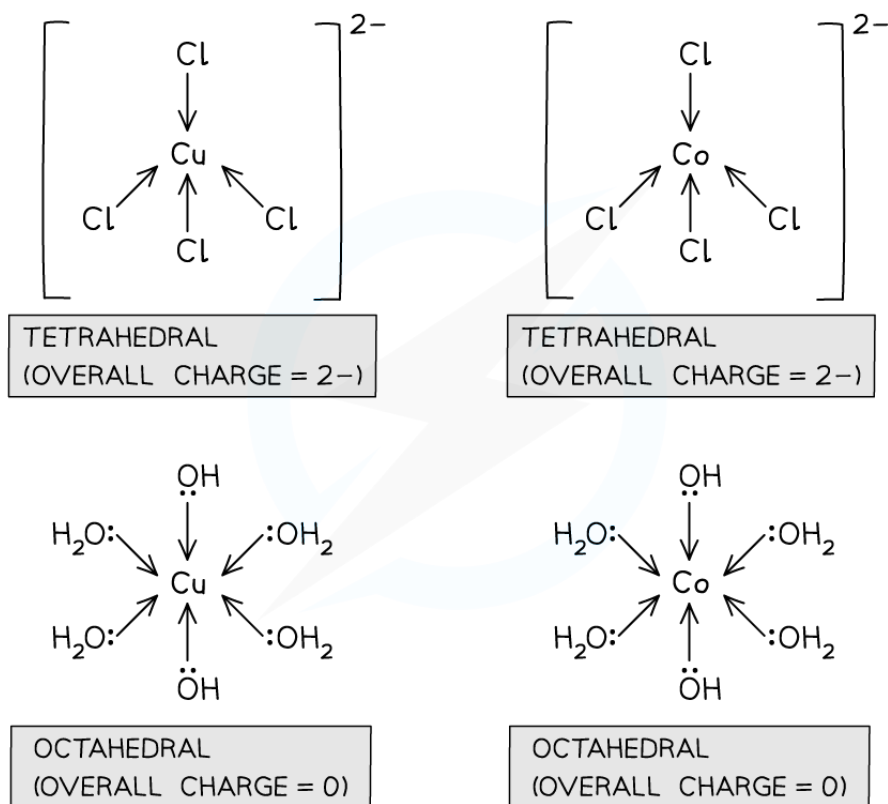
Complexes with hydroxide & chloride ions

- Hydroxide and chloride ions are examples of **negatively** charged ligands
- Both ligands contain a lone pair of electrons which can be used to form a dative covalent bond with the central metal ion
- Hydroxide ligands are small, so 6 of them can fit around a central metal ion and the complex formed will have a **coordination number** of 6
- Chloride ligands are large ligands, so only 4 of them will fit around a central metal ion
- Complexes with 4 chloride ligands will have a **coordination number** of 4
 - Complexes with 4 chloride ligands will form **tetrahedral** complexes
 - Whereas hydroxide ligands will form **octahedral** complexes
- A complex with cobalt(II) or copper(II) as a central metal ion and chloride ions as ligands, will have an **overall charge** of 2-
 - The central metal ion has a charge of 2+
 - Each chloride ligand has a charge of 1-
 - There are 4 chloride ligands in the complex, so the overall **negative** charge is 4-



Your notes

- The overall **positive** charge is 2+
- Therefore, the overall charge of the complex is 2-
- A complex with cobalt(II) or copper(II) as a central metal ion and hydroxide ions as ligands, will have no overall charge
 - The central metal ion has a charge of 2+
 - Each hydroxide ligand has a charge of 1-
 - There are 2 hydroxide ligands in the complex, so the overall **negative** charge is 2-
 - The overall **positive** charge is 2+
 - Therefore, the overall charge on the complex is 0



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Cobalt(II) and copper(II) form tetrahedral complexes with chloride and octahedral complexes with water and hydroxide ligands



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6.2.4 Ligands

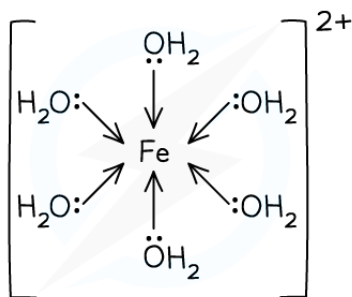
Define Ligand

- A **ligand** is a molecule or ion that has **one** or **more** lone pairs of electrons
- These lone pairs of electrons are donated by the ligand, to form **dative covalent bonds** to a central metal atom or ion

Examples of ligands table

Name of Ligand	Formula of Ligand
Water	H_2O
Ammonia	NH_3
Chloride	Cl^-
Cyanide	CN^-
Thiocyanate	SCN^-
Ethanedioate (ox)	$^-\text{COO}-\text{COO}^-$ ($\text{C}_2\text{O}_4^{2-}$)
1,2-diaminoethane (en)	$\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$

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Example of a complex formed between a transition metal ion (Fe^{2+}) and a ligand (H_2O) by dative covalent bonds



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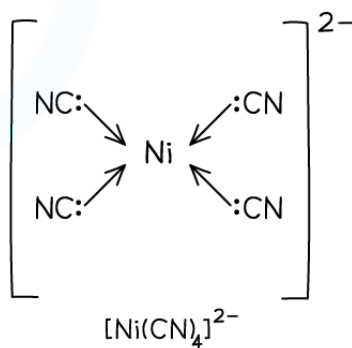
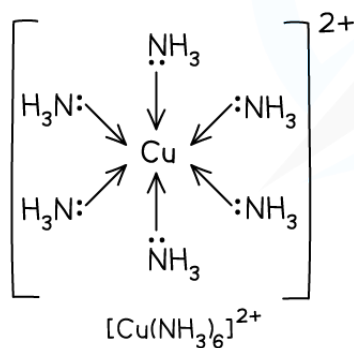
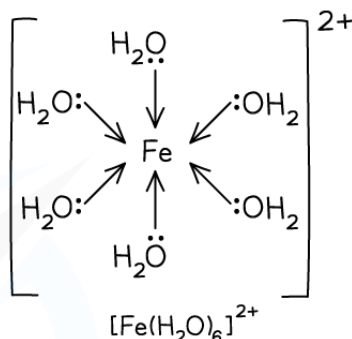
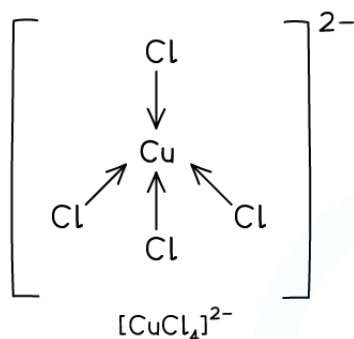
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Types of Ligands

- Different **ligands** can form different numbers of dative bonds to the central metal ion in a complex.
 - Some ligands can form **one** dative bond to the central metal ion
 - Other ligands can form **two** dative bonds, and some can form **multiple** dative bonds

Monodentate ligands

- Monodentate** ligands can form only **one** dative bond to the central metal ion
- Examples of monodentate ligands are:
 - Water (H_2O) molecules
 - Ammonia (NH_3) molecules
 - Chloride (Cl^-) ions
 - Cyanide (CN^-) ions



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Examples of complexes with monodentate ligands

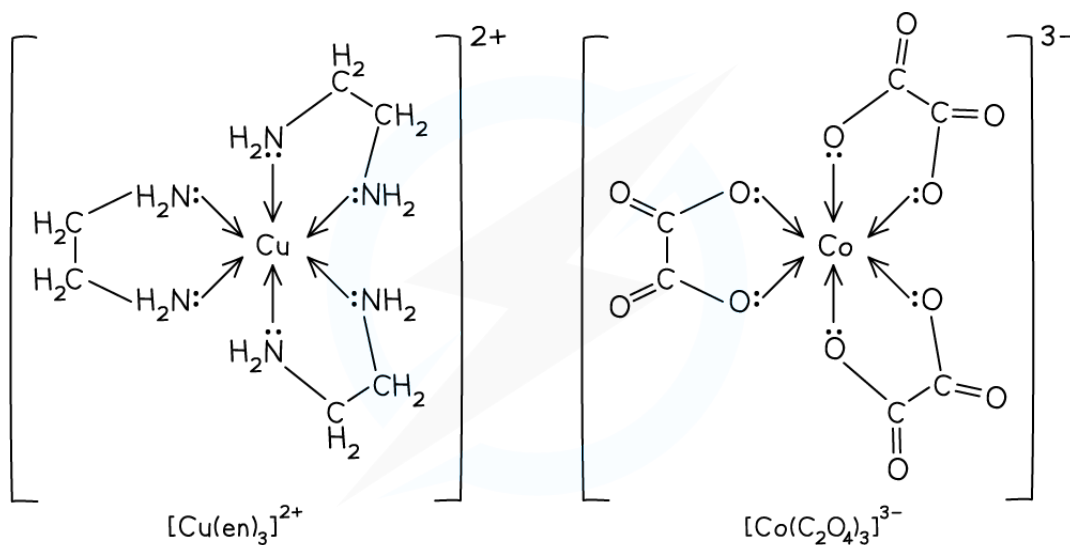
Bidentate ligands

- Bidentate** ligands can each form **two** dative bonds to the central metal ion

- This is because each ligand contains **two** atoms with lone pairs of electrons
- Examples of bidentate ligands are:
 - 1,2-diaminoethane ($\text{H}_2\text{NCH}_2\text{CH}_2\text{NH}_2$) which is also written as '**en**'
 - Ethanedioate ion ($\text{C}_2\text{O}_4^{2-}$) which is sometimes written as '**ox**'



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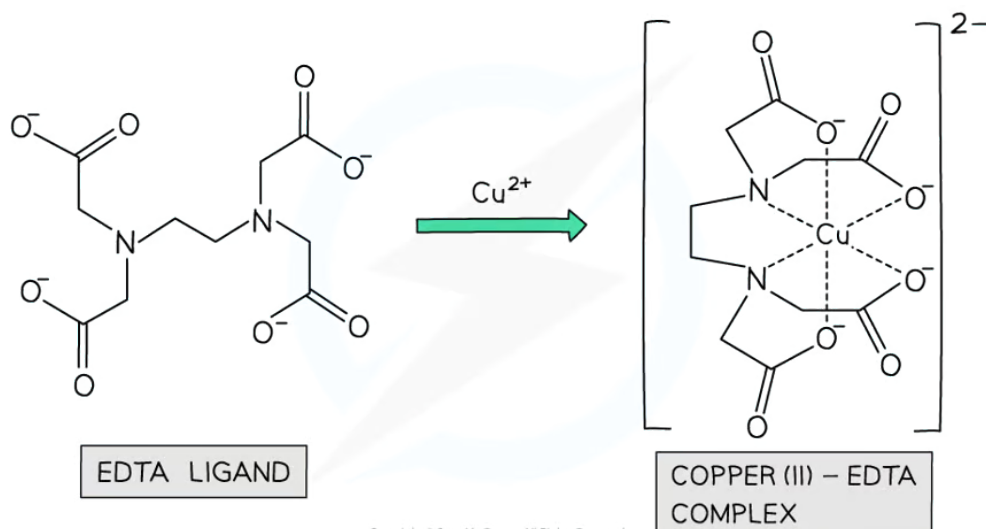


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Examples of complexes with bidentate ligands

Polydentate ligands

- Some ligands contain more than two atoms with lone pairs of electrons
- These ligands can form more than two dative bonds to the central metal atom and are said to be **polydentate** ligands
- An example of a polydentate ligand is EDTA^{4-} , which is a **hexadentate** ligand as it forms 6 dative covalent bonds to the central metal ion



Example of a polydentate ligand complex

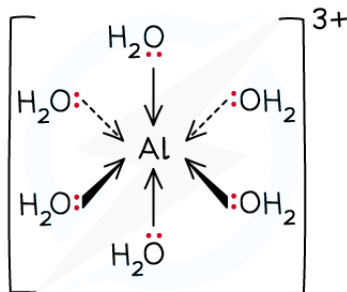
6.2.5 Defining Complexes



Your notes

Define Complex

- A **complex** is a molecule or ion formed by a **central metal atom or ion** surrounded by one or more **ligands**
- A ligand is a species that contains one or more **lone pairs of electrons**
- The ligand forms a **dative covalent bond** with the central metal atom or ion, by donating its lone pair of electrons



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Example of a complex ion formed by a central aluminium(III) ion and six water molecule ligands



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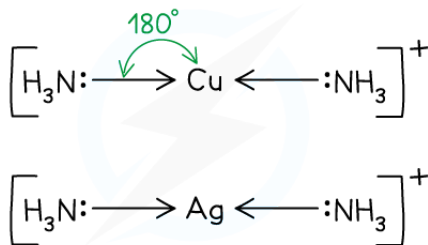
6.2.6 Geometry of Complexes

Geometry of the Transition Element Complexes

- Depending on the **size of the ligands** and the **number of dative bonds** to the central metal ion, transition element complexes have different geometries
 - Dative bonds can also be referred to as **coordinate bonds**, especially when discussing the geometry of a complex

Linear

- Central metal atoms or ions with **two coordinate bonds** form **linear** complexes
- The bond angles in these complexes are 180°
- The most common examples are a copper (I) ion, (Cu^+), or a silver (I) ion, (Ag^+), as the central metal ion with two coordinate bonds formed to two ammonia ligands

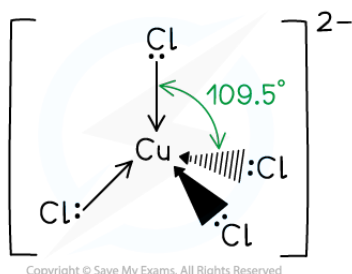


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Example of a linear complex

Tetrahedral

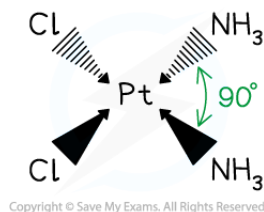
- When there are **four coordinate bonds** the complexes often have a **tetrahedral** shape
 - Complexes with four **chloride ions** most commonly adopt this geometry
 - Chloride ligands are large, so only four will fit around the central metal ion
- The bond angles in tetrahedral complexes are 109.5°



Example of a tetrahedral complex

Square planar

- Sometimes, complexes with **four coordinate bonds** may adopt a **square planar** geometry instead of a tetrahedral one
 - Cyanide ions (CN^-) are the most common ligands to adopt this geometry
 - An example of a square planar complex is **cisplatin**
- The bond angles in a square planar complex are 90°



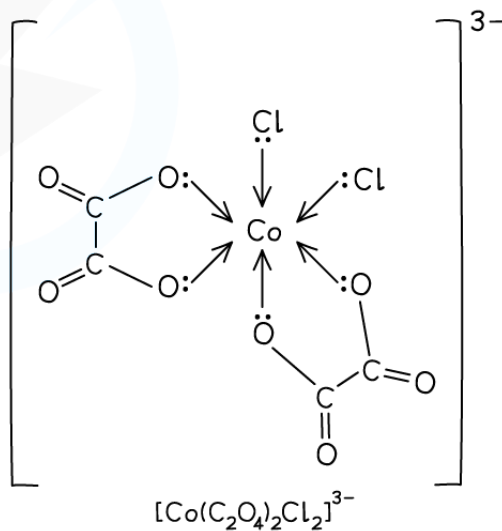
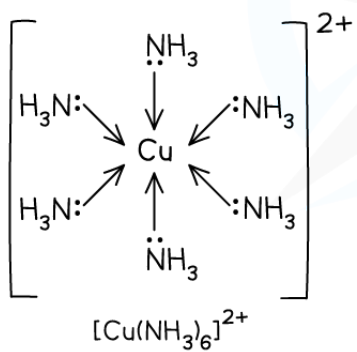
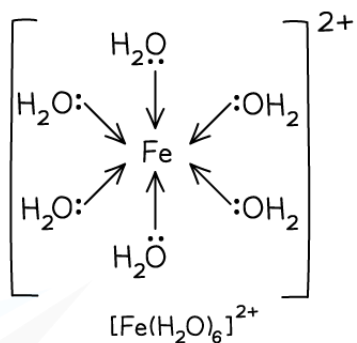
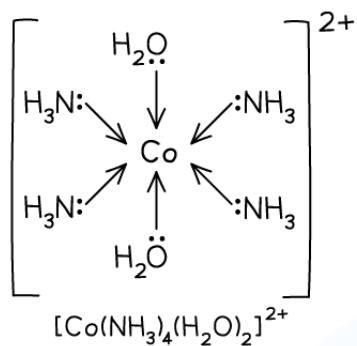
Cisplatin is an example of a square planar complex

Octahedral

- Octahedral** complexes are formed when a central metal atom or ion forms **six coordinate bonds**
- This could be six coordinate bonds with **six** small, **monodentate** ligands
 - Examples of such ligands are **water** and **ammonia** molecules and **hydroxide** and **thiocyanate** ions
- It could be six coordinate bonds with **three bidentate** ligands
 - Each bidentate ligand will form two coordinate bonds, meaning six coordinate bonds in total
 - Examples of these ligands are **1,2-diaminoethane** and the **ethanedioate ion**
- It could be six coordinate bonds with **one polydentate** ligand
 - The polydentate ligand, for example EDTA^{4-} , forms all six coordinate bonds
- The bond angles in an octahedral complex are 90°



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Examples of octahedral complexes

Types of ligands table



Your notes

Geometry	Number of Coordinate (Dative) Bonds	Bond Angle (°)	Ligand	Formula
Linear	2	180	Ammonia	NH ₃
Tetrahedral	4	109.5	Chloride ion	Cl ⁻
Square planar	4	90	Cyanide ion	CN ⁻
Octahedral	6	90	Water Ammonia Hydroxide ion Thiocyanate ion Ethanedioate ion 1,2-diaminoethane EDTA ⁴⁻	H ₂ O NH ₃ OH ⁻ SCN ⁻ ⁻ OOC-COO ⁻ (C ₂ O ₄ ²⁻) NH ₂ CH ₂ CH ₂ NH ₂

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Coordination Number & Predicting Complex Ion Formula & Charge

- The **coordination number** of a complex is the number of coordinate bonds that are formed between the **ligand(s)** and the central metal atom or ion
- Some ligands can form only one coordinate bond with the central metal ion (**monodentate ligands**), whereas others can form two (**bidentate ligands**) or more (**polydentate ligands**)
- It is **not** the number of ligands which determines the coordination number, it is the number of coordinate (dative) bonds

Predicting complex ion formula & charge

- The formula and charge of a complex ion can be predicted if the following are known:
 - The central metal ion and its charge/oxidation state
 - The ligands
 - The coordination number/geometry

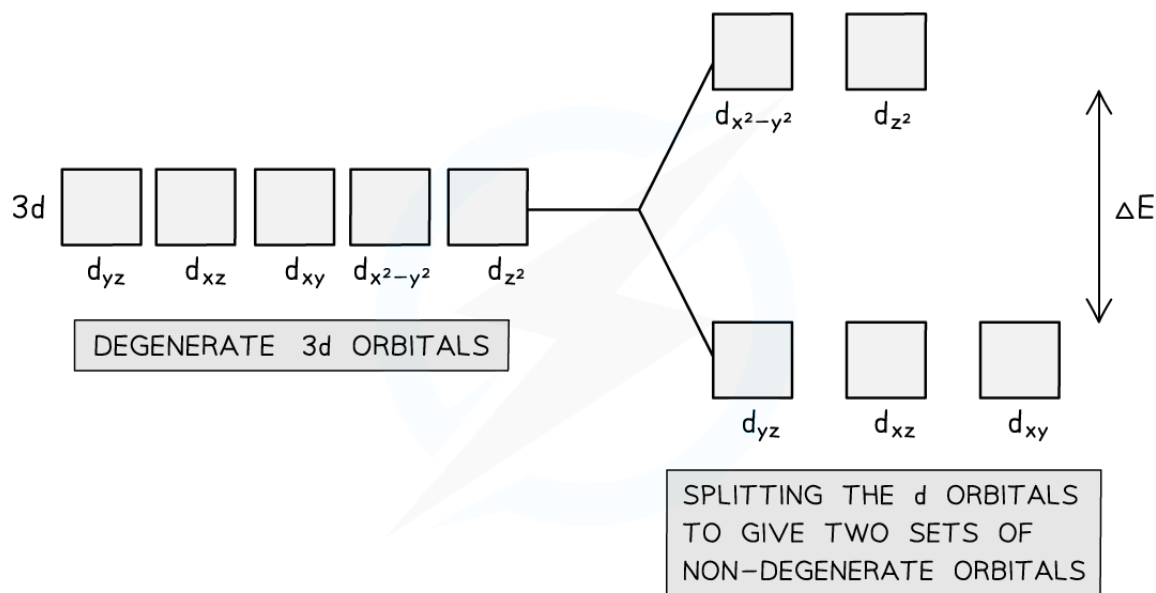


Your notes

6.2.7 Degenerate & non-Degenerate d Orbitals

Define Degenerate & non-Degenerate d Orbitals

- There are **five d orbitals** in an **isolated** transition element atom or ion
 - An isolated transition element is one that is not bonded to anything else
- These d orbitals are all at the same energy level (they are equal in energy) and are therefore said to be **degenerate orbitals**
- When ligands are attached, the transition element ion is **not** isolated anymore
- The dative bonding from the ligands causes the five d orbitals to split into two sets
- These two sets are **not** equal in energy and are described as being **non-degenerate orbitals**



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Upon binding to ligands, the d orbitals of the transition element ion split into two non-degenerate sets of orbitals



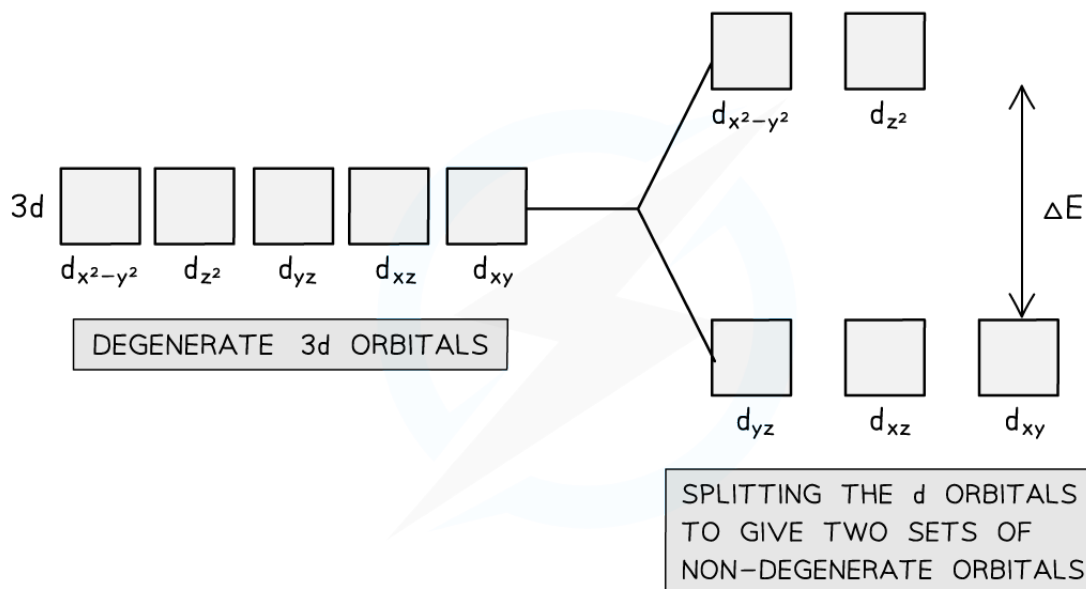
Your notes

Degenerate d Orbital Splitting

- An isolated transition element has five **degenerate** 3d orbitals
- Upon dative covalently bonding to a ligand, these d orbitals are split into two sets of non-degenerate orbitals

Splitting in octahedral complexes

- In **octahedral** complexes, there are six ligands arranged around the **central metal ion**
- The lone pairs of the ligands repel the electrons in the x^2-y^2 and z^2 orbitals of the metal ion **more than** they repel the electrons in the $3d_{yz}$, $3d_{xz}$, and $3d_{xy}$ orbitals
- This is because the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals line up with the dative bonds in the complex's octahedral shape
- This is because the ligands are attached to or approaching the central metal ion along the x, y and z axes, and the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals have lobes along these axes
- The electrons in these two orbitals are closer to the bonding electrons, so there is more **repulsion**
- This means that when the d orbitals split, the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals are at a slightly higher energy level than the other three
- The difference in energy between the **non-degenerate** d orbitals is labelled as ΔE



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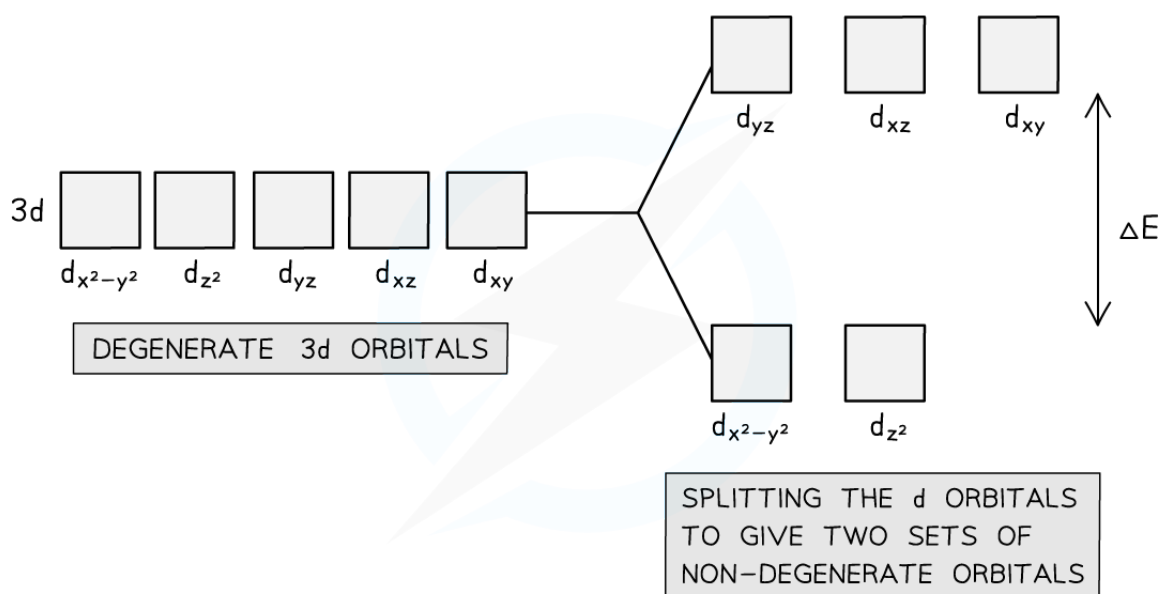
Splitting of 3d orbitals in an octahedral complex

Splitting in tetrahedral complexes



Your notes

- In **tetrahedral** complexes, there are four ligands arranged around the central metal ion
- The bonding pair of electrons from the four ligands now line up with the $3d_{yz}$, $3d_{xz}$, and $3d_{xy}$ orbitals of the central metal ion
- Now, the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals lie between the metal-ligand bonds
- Therefore, there is less repulsion with the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals
- When the d orbitals split this time, the $3d_{x^2-y^2}$ and $3d_{z^2}$ orbitals are at **lower** and **more stable** energy level than the other three



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Splitting of 3d orbitals in a tetrahedral complex



Your notes

6.2.8 Coloured Complexes

Coloured Compounds & Electron Promotion

- Most transition element complexes are **coloured**
- A transition element complex solution which is coloured, **absorbs** part of the electromagnetic spectrum in the visible light region
- The observed colour is the **complementary colour** which is made up of light with frequencies that are **not** absorbed
 - For example, copper(II) ions absorb light from the **red** end of the spectrum
 - The complementary colour observed is therefore **pale blue** (cyan)



The visible light region of the electromagnetic spectrum

Electron promotion

- In an **isolated transition element ion** (which is not bonded to any ligands), all of the 3d orbitals are **degenerate**
- However, when ligands are attached to the central metal ion through dative covalent bonds, these orbitals are split into two sets of **non-degenerate orbitals**
- The difference in energy between these two sets of orbitals is ΔE
- When light shines on a solution containing a transition element complex, an electron will absorb this exact amount of energy (ΔE)
- The amount of energy absorbed can be worked out by the equation:

$$\Delta E = h \times \nu$$

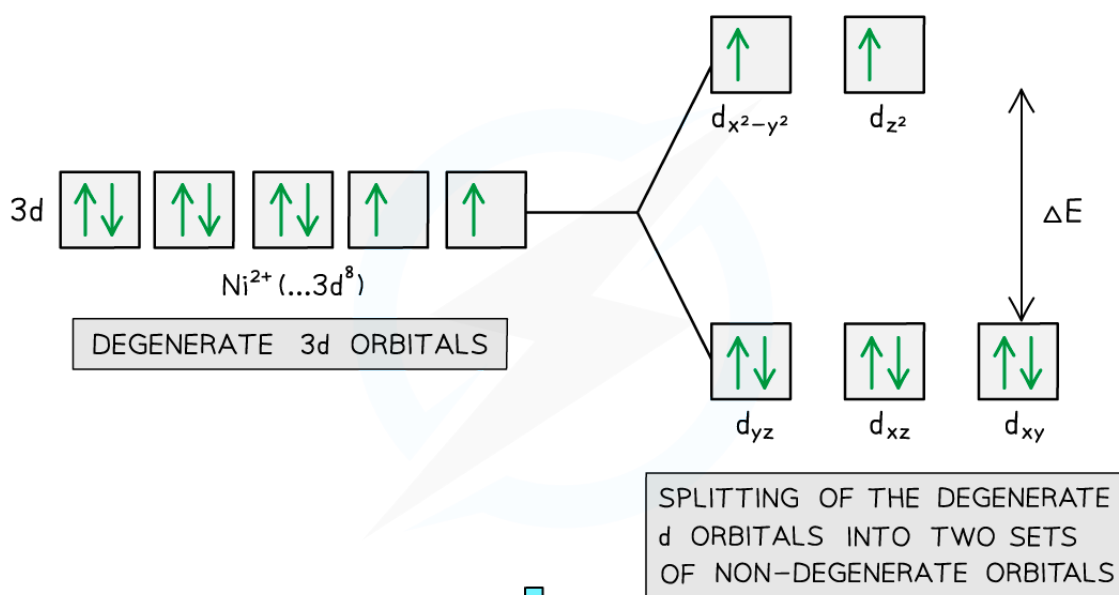
h = Planck's constant ($6.626 \times 10^{-34} \text{ m}^2 \text{ kg s}^{-1}$)

ν = frequency (Hertz, Hz or s^{-1})



Your notes

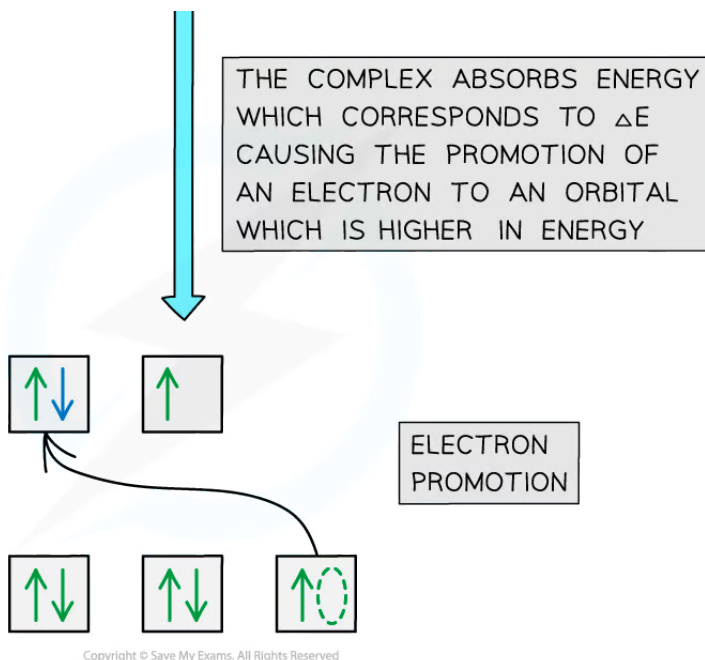
- The electron uses the energy from the light to jump into a higher, non-degenerate energy level
 - This is also called **electron promotion**
- The other frequencies of light which are **not absorbed** combine to make the **complementary colour**
- The diagram below shows an example of electron promotion in an octahedral complex of a nickel(II) Ni^{2+} ion



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Your notes



Electron promotion in a Ni(II) complex when light shines on the solution

Effects of Ligands on Complementary Colour

- Transition element complexes absorb the frequency of light which corresponds to the exact energy difference (ΔE) between their **non-degenerate** d orbitals
- The frequencies of light which are not absorbed combine to make the **complementary colour** of the complex
- It is the complementary colour which is seen
- However, the exact energy difference (ΔE) is affected by the different ligands which surround the transition element ion
- Different ligands will split the d orbital by a different amount of energy
- This depends on the **repulsion** that the d orbital experiences from these ligands
- Therefore, the **size** of ΔE and thus the **frequency** of light absorbed by the electrons will be slightly different
- As a result, a different colour of light is absorbed by the complex solution and a different **complementary colour** is observed
- This means that complexes with **similar transition elements ions**, but **different ligands**, can have different colours
 - For example, the $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ complex has a **light blue** colour
 - Whereas the $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ has a **dark blue** colour
 - Despite the copper ion having an oxidation state of +2 in both complexes
 - This is evidence that the ligands surrounding the **complex ion** affect the colour of the complex



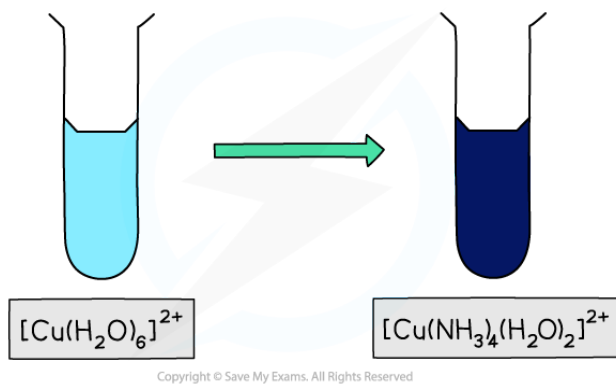
Your notes

Ligand Exchange in Copper(II) & Cobalt(II) Complexes

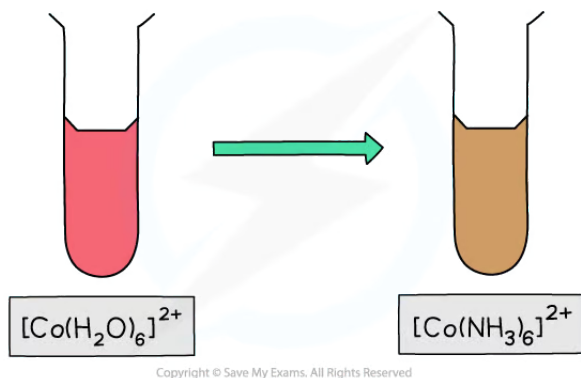
- Different ligands may affect the complementary colour of a transition ion complex solution
- This is shown by ligand exchange reactions in copper(II) and cobalt(II) complexes, as this causes a change in colour of the complexes

Copper(II) & cobalt(II) ions

- The ligand exchange of $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ and $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ by NH_3 ligands causes a change in the colour of the solutions
 - $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ is **light blue** in colour whereas $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ is **deep blue** in colour
 - $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ is a **pink** solution whereas $[\text{Co}(\text{NH}_3)_6]^{2+}$ is a **brown** solution
- The colour change results from the ammonia ligands, which cause the d orbitals to split by a different amount of energy (ΔE)
- Therefore, the size of ΔE and the frequency of light absorbed by the electrons will be slightly different
- As a result, a different colour of light is absorbed and thus a different **complementary colour** is observed

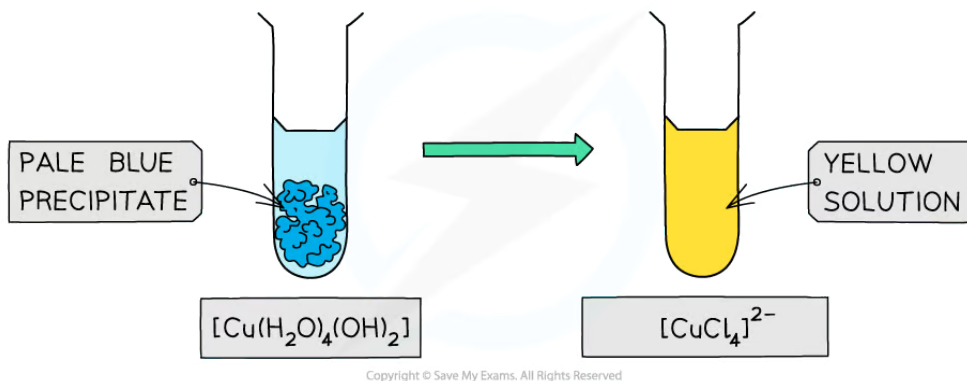
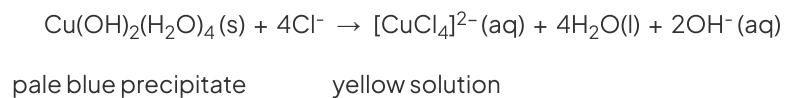


Ligand exchange of the water ligands by ammonia ligands causes a change in colour of the copper(II) complex solution

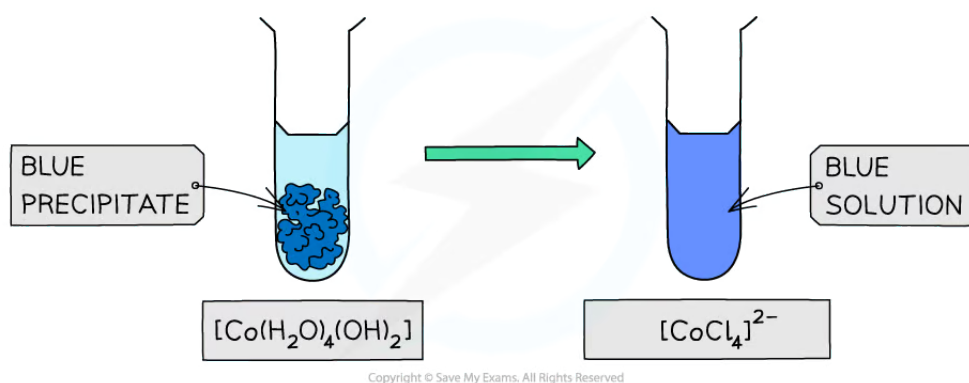


Ligand exchange of the water ligands by ammonia ligand causes a change in colour of the cobalt(II) complex solution

- Similarly, full ligand exchange by chloride ions in copper(II) and cobalt(II) complexes results in a change in complementary colour



Ligand exchange by chloride ligands causes a change in colour of the copper(II) complex solution



Ligand exchange by chloride ligands causes a change in colour of the cobalt(II) complex solution

- As before, this suggests that different ligands will split the d orbitals differently